

#### 4.1 the gaseous state: ideal and real gases and $pV = nRT$

Pressure in a gas is due to the fact that the gas molecules in the container have a mass and velocity, which means they have a momentum. When the particles collide with the walls of the container, their direction and velocity changes and so does their momentum. Rate of change of momentum is force, this force is being applied over an area. Pressure = force over area

##### Ideal gases according to kinetic theory:

made up of molecules which are in constant random motion in straight lines

Molecules behave as rigid spheres

Pressure is due to collisions between the molecules and the walls of the container

All collisions between the molecules themselves and between molecules and walls of the container are perfectly elastic which means there's no loss of kinetic energy during the collision

The temperature of the gas is proportional to the average kinetic energy of the molecules

##### Relationship between pressure, volume and temperature:

$$\text{pressure} \times \text{volume} = \text{moles} \times \text{gas constant} \times \text{temperature}$$

**Pressure** has to be in pascals which is equivalent to  $\text{Nm}^{-2}$

**Volume** has to be in  $\text{m}^3$

**Moles** are in mole

**Gas constant** is  $8.31 \text{ JK}^{-1}\text{mol}^{-1}$

**Temperature** is in kelvin

##### Dalton's law:

$$PT = p_a + p_b$$

When asked to find the relative formula mass of a gas from its density, there may be errors as no gas is actually ideal, and the density used may have not been accurate

##### Perfect conditions for ideal gases:

when **pressures** are low and 100kPa or 1 atm as molecules are more spread apart so they are less likely to be forced together and compelled to form attraction

when **temperature** is high as particles move faster and quicker thus there's less and less chance of intermolecular interactions

##### The most ideal gas:

Has the smallest possible molecules and lowest possible IM forces, which is helium

#### 4.2 Bonding and structure

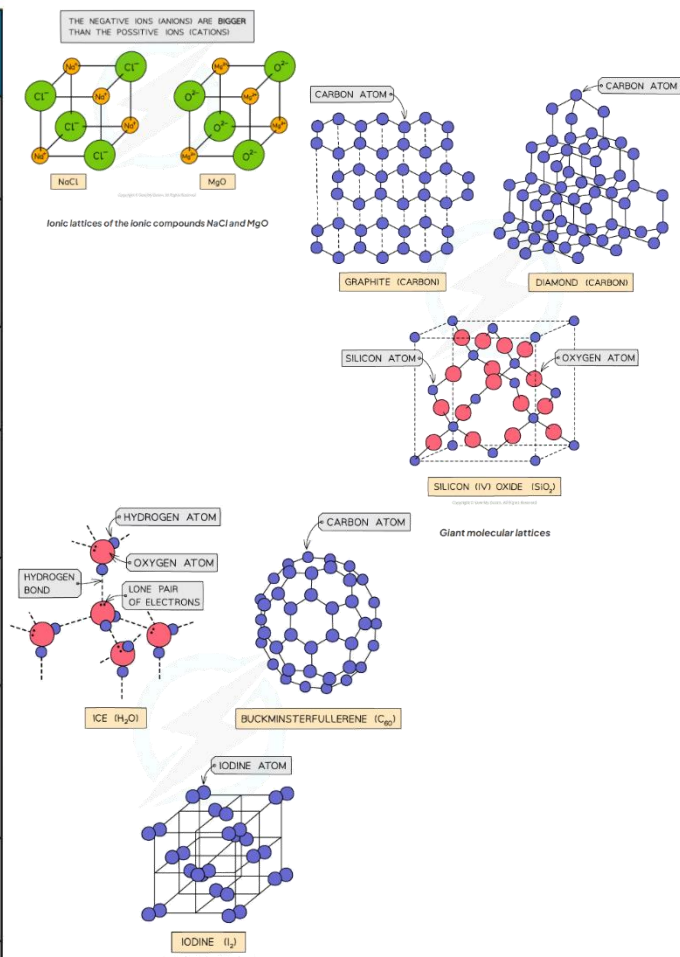
**Giant ionic:** due to electrostatic forces between a positively charged cation and a negatively charged anion, the ion compounds are arranged in alternating lattices, they are brittle as the ionic crystals can split apart, they have high melting and boiling point, they're both increase with charge density of the ions. They are soluble in water as they can form ion-permanent dipole-dipole bonds, they can conduct electricity when molten or in solution as they have free ions

**Simple molecular:** bonds where nonmetals share electrons, they have low melting and boiling points as they have weak IM forces between their molecules, they are insoluble in water unless they are polar or can form hydrogen bonds, they cannot conduct electricity in any state, except for a few examples in solution such as HCl

**Giant molecular:** they have high melting and boiling points as there are lots of covalent bonds linking the whole structure, they can be hard or soft such as graphite or diamond respectively, mostly insoluble in water and mostly do not conduct electricity

**Giant metallic:** due to electrostatic forces between positively charged cations and negatively charged electrons, often packed in hexagonal layers or in a cubic arrangement, they are malleable as layers can slide over each other, the attractive forces between metal ions and electrons act in all directions, the lattice cannot break and only changes shape. Metals are strong and hard due to the strong attractive forces between the metal ions and the delocalized electrons, they have high melting and boiling points, they can conduct electricity when solid or liquid as they have free electrons

	Giant Ionic	Giant Metallic	Simple Covalent	Giant Covalent
<b>Melting and Boiling Points</b>	High	Moderately high to high	Low	Very high
<b>Electrical Conductivity</b>	Only when molten or in solution	When solid or liquid	Do not conduct electricity	Do not conduct electricity (except for graphite)
<b>Solubility</b>	Soluble	Insoluble but some may react	Usually insoluble unless they are polar	Insoluble
<b>Hardness</b>	Hard, brittle	Hard, malleable	Soft	Very hard (diamond and $\text{SiO}_2$ ) or soft (graphite)
<b>Physical State at Room Temperature</b>	Solid	Solid	Solid, liquid or gas	Solid
<b>Forces</b>	Electrostatic attraction between ions	Delocalised sea of electrons attracting positive ions	Weak intermolecular forces between molecules and covalent bonds within a molecule	Electrons in covalent bonds between atoms
<b>Particles</b>	Ions	Positive ions in a sea of electrons	Small molecules	Atoms
<b>Examples</b>	NaCl	Copper	$\text{Br}_2$	Graphite, silicon(IV) oxide



Property	Diamond	Graphite	Silicon (IV) Oxide
<b>Melting point</b>	Very high as many covalent bonds must be broken to separate atoms	Very high as many covalent bonds must be broken to separate atoms	Very high as many covalent bonds must be broken to separate atoms
<b>conductivity</b>	Non-conductor as all 4 carbon electrons are involved in bonding	Conductor as only three carbon electrons are used for bonding	Non-conductor as it has no free electrons
<b>strength</b>	Strong, each carbon is joined to 4 others rigidly	Soft as each carbon is joined to 3 others in a layered structure held by van der Waals allowing them to slide over each other	Strong as each silicon atom is joined to 4 oxygen, and each oxygen is joined to 2 silicon atoms
<b>solubility</b>	Insoluble in water and organic solvents as no possible attractions could occur between them	Insoluble in water and organic solvents as no possible attractions could occur between them	Insoluble in water and organic solvents as no possible attractions could occur between them

**Allotropes:** two or more different forms of an element in the same physical state, having different atomic arrangements such as buckminsterfullerene, graphite and diamond

